

# Chapter-5

## Lesson-1

### I – The Nature of Energy

- A **system** is the defined part of the universe under study.  
A system could be a body or a chemical reaction.
- Everything outside of the system will be called the **surroundings**.
- **Thermochemistry** is the branch of thermodynamics that focuses on the heat involved in chemical reactions.

#### A) Kinetic Energy and Potential Energy

- **Energy** is the capacity to move a body against a force.
- **Kinetic energy** is the energy a body has due to its motion.

$$E_k = \frac{1}{2} m v^2$$

where  $E_k$  is the kinetic energy of the body (J),  $m$  is the mass of the body (kg), and  $v$  is the velocity of the body (m/s).

- **Potential energy** is the energy a body has due to its relative position to another body.

$$E_p = mgh$$

where  $E_p$  is the gravitational potential energy of a body (J),  $m$  is the mass of the body (kg),  $g$  is earth's gravitation constant [ $9.8 \text{ kg}\cdot\text{m/s}^2$ ], and  $h$  is the height of the body above some point on earth (m).

$$E_{el} = \frac{kQ_1Q_2}{d} \quad \text{Coulombic Energy}$$

where  $E_{el}$  is the electrostatic potential energy (J) between two charged bodies,  $k$  is a constant [ $8.99 \times 10^9 \text{ J}\cdot\text{m/C}^2$ ],  $Q_1$  and  $Q_2$  are the electrical charges on the two bodies, and  $d$  is the distance between them.

$$F = \frac{kQ_1Q_2}{d^2} \quad \text{Coulombic Force}$$

where  $F$  is the attractive or repulsive force between two charged bodies (J/m).

**Q1:** What is the kinetic energy of one He-4 atom moving at 1360m/s?

**A1:**  $E_k = \frac{1}{2} mv^2$

$$E_k = \frac{1}{2}(0.00400\text{kg/mol} \div 6.022 \times 10^{23} \text{mol})(1360\text{m/s})^2 = 6.14 \times 10^{-21} \text{ J}$$

**Q2:** What is the potential energy between two repelling electrons ( $1.602 \times 10^{-19}\text{C}$ ) that are 148 pm apart?

**A2:**  $E_{el} = \frac{kQ_1Q_2}{d}$

$$E_p = \frac{(8.99 \times 10^9 \text{ J}\cdot\text{m/C}^2)(-1.602 \times 10^{-19}\text{C})^2}{148 \times 10^{-12} \text{ m}} = 1.56 \times 10^{-18} \text{ J}$$

**Q3:** What is the force of repulsion between two electrons that are 148 pm apart?

**A3:**  $F = \frac{kQ_1Q_2}{d^2}$

$$F = \frac{(8.99 \times 10^9 \text{ J}\cdot\text{m/C}^2)(-1.602 \times 10^{-19}\text{C})^2}{(148 \times 10^{-12} \text{ m})^2} = 1.05 \times 10^{-8} \text{ J/m}$$

## B) Transferring Energy: Heat and Work

- Energy is transferred between systems and surroundings in two general ways, as heat or as work.
- **Heat** ( $q$ ) is the form of energy that flows between two systems at different temperatures. Heat will flow spontaneously from the system at the higher temperature to the system at the lower temperature until the two bodies reach the same temperature.

$$q = mc\Delta T$$

where  $q$  is heat (J),  $m$  is mass (g),  $c$  is specific heat capacity ( $\text{J/g}\cdot^\circ\text{C}$ ), and  $\Delta T$  is change in temperature ( $^\circ\text{C}$ ).

- **Work** ( $w$ ) is done when matter is moved against a force.

$$w = Fd(\cos\Phi)$$

where  $w$  is work (J),  $F$  is force (J/m),  $d$  is distance moved (m), and  $\Phi$  is the angle between the force and displacement vectors.

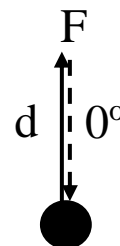
**Q4:** How much work is done if a 5.00 kg (11.0 lbs) stone is lifted vertically 1.000 m? *When lifting a body vertically,  $F = mg$ .*

**A4:**  $w = Fd(\cos\Phi) = (m)(g)(d)(\cos\Phi)$  where “g” is gravity

$$w = (5.00 \text{ kg})(9.8 \text{ kg}\cdot\text{m/s}^2)(1.000 \text{ m})(\cos 0)$$

$$w = \boxed{+49.0 \text{ J}}$$

earth's gravity



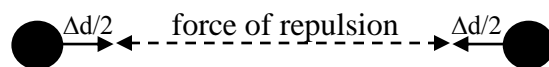
- **Work** (w) is done when charged matter is moved against a force.

$$w = -kQ_1Q_2 \left[ \frac{1}{d_f} - \frac{1}{d_i} \right]$$

where **w** (J) is work, **k** is a constant [ $8.99 \times 10^9 \text{ J}\cdot\text{m/C}^2$ ], **Q<sub>1</sub>** and **Q<sub>2</sub>** are the electrical charges on the two charged bodies (C), and **d** is the distance moved (m).

**Q5:** How much work is done if two electrons are moved 1 pm closer from a distance of 148 pm to 147 pm?

**A5:**  $w = -kQ_1Q_2 \left[ \frac{1}{d_f} - \frac{1}{d_i} \right]$



$$w = - (8.99 \times 10^9 \text{ J}\cdot\text{m/C}^2)(-1.602 \times 10^{-19} \text{ C})^2 [-4.60 \times 10^7 \text{ m}]$$

$$w = \boxed{+1.06 \times 10^{-20} \text{ J}}$$

## II – The First Law of Thermodynamics

- **Thermodynamics** is the study of energy and its interconversions.
- **The first law of thermodynamics** states that energy can neither be created nor destroyed but can change form. Thus the total energy of the universe is always conserved!

$$\Delta E_{\text{uni}} = 0$$

- Notice that the first law of thermodynamics is the law of conservation of energy.
- During a chemical change in a system, the energy change of the system is equal to but opposite in sign to the energy change of the surroundings.

$$\Delta E_{\text{sys}} + \Delta E_{\text{sur}} = 0$$

## A) Internal Energy (E):

- The **internal energy** (E) of a system is the sum of all the kinetic and potential energy present in the system.
- Internal energy (E) includes such things as:
  - attractive energy between protons & electrons
  - bond energy between atoms
  - attractive energy between molecules
  - kinetic energy of particles (*translational, rotational, & vibrational*)
- Internal Energy (E) cannot be determined. However, the change in internal energy ( $\Delta E$ ) can be determined.
- In any process, the change in internal energy of a system ( $\Delta E$ ) is equal to the sum of the heat (q) exchanged by the system and the surroundings plus any work (w) done between the system and the surroundings.

$$\Delta E = q + w$$

q is (+) if heat is absorbed by the system from the surroundings.

q is (−) if heat is released by the system to the surroundings.

w is (+) if work is done on the system by the surroundings.

w is (−) if work is done by the system on the surroundings.

**Q6:** Suppose a sample of gas absorbs 50.0 kJ of heat, expands and does 20.0 kJ of work on the surroundings. What is the energy change of the gas system?

**A6:**  $\Delta E = q + w$

$$\Delta E = (+50.0 \text{ kJ}) + (-20.0 \text{ kJ})$$

$$\Delta E = \boxed{+30.0 \text{ kJ}}$$

**Q7:** If a system did 65.0 kJ of work as its internal energy decreased by 90.0 kJ, then how much heat did it exchange with its surroundings?

**A7:**  $\Delta E = q + w$

$$-90.0 \text{ kJ} = (q) + (-65.0 \text{ kJ})$$

$$q = \boxed{-25.0 \text{ kJ}}$$

**Q8:** In an aquarium, a battery drives the electric motor of the water pump. Together, the battery and electric motor do 555 kJ of work on the pump while releasing 124 kJ of heat. What is the change in the internal energy of the battery and electric motor?

**A8:**  $\Delta E = q + w$

$$\Delta E = (-124 \text{ kJ}) + (-555 \text{ kJ})$$

$$\Delta E = \boxed{-679 \text{ kJ}}$$

**Q9:** The reaction for the combustion of methane in a piston chamber is:  $\text{CH}_{4(g)} + 2 \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2 \text{H}_2\text{O}_{(g)} + 890.0 \text{ kJ}$ . If the reaction mixture loses 400.0 kJ of heat, then what is the change in work of the system?

**A9:**  $\Delta E = q + w$

$$-890.0 \text{ kJ} = (-400. \text{ kJ}) + (X)$$

$$X = \boxed{-490.0 \text{ kJ}}$$

## B) Endothermic and Exothermic Reactions:

- In an **endothermic reaction**, the reaction (system) absorbs heat from the surroundings (and then stores the energy).

Thus the products of the reaction (system) have more potential (stored) energy than before.

In an endothermic reaction, the surroundings lose (release) heat and there is usually a decrease in the surrounding's temperature.

- In an **exothermic reaction**, the reaction (system) releases heat (that was stored energy) to the surroundings.

Thus the products of the reaction (system) have less potential (stored) energy than before.

In an exothermic reaction, the surroundings gain (absorb) heat and there is usually an increase in the surrounding's temperature.

### C) State Functions:

- A **state function** is a property of a system that depends only on the present state (ie: composition, volume, temperature, pressure, location, etc.) of the system and not on the path the system took to reach that state.
- Internal energy (E) is an example of a state function. However, heat (q) and work (w) are not state functions.
- A change in state function depends only on its initial and final states and not on how the change took place.

$$\Delta E = E_f - E_i$$

- Consider the combustion of octane in an open barrel and in a car engine.



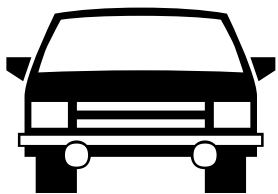
The energy change ( $\Delta E$ ) is the same in both (10943 kJ released), but the amount of heat released and work done are very different.

In the barrel, the  $\Delta E$  is due to almost 100% to heat released.



$$\Delta E = q + w$$

In the engine the  $\Delta E$  is due to both heat released and work done.



$$\Delta E = q + w$$

Since heat (q) and work (w) are not the same in both cases, then heat and work are not state functions! ***For heat and work, how the octane burns makes a difference to their values!***

Examples of state functions other than internal energy (E) are enthalpy (H), entropy (S), free energy (G), temperature (T), pressure (P), and volume (V).

# Chapter-5

## Lesson-2

### III - Enthalpy

- A system's enthalpy (H) is equal to its internal energy + the product of the system's pressure and volume.

$$H = E + PV$$

- Enthalpy is a state function because internal energy, pressure, and volume are all state functions.
- In a system open to the atmosphere, the only form of work done by a chemical or physical change is the mechanical work associated with a change in the system's volume.

$$w = - P_{\text{external}}\Delta V$$

where **P** is the atmospheric pressure and  $\Delta V$  is the system's volume change. This form of work is called PV work.

- In chemical reactions that occur at constant pressure, **PV** work is often due, in part, to the change in the moles of gas. If more moles of gas is produced than reacted, then the volume of the system will expand and **PV** work will be equal to  $\Delta n(RT)$ .
- The change in enthalpy of a system open the atmosphere is:

$$\Delta H = \Delta E + P_{\text{external}}\Delta V$$

Since

$$\Delta E = q + w$$

then

$$\Delta H = q + w + P_{\text{external}}\Delta V$$

and since

$$w = - P_{\text{external}}\Delta V$$

then

$$\Delta H = q + \cancel{w} + P_{\text{external}}\Delta V$$

or

$$\Delta H = q_p \quad (\text{at constant pressure})$$

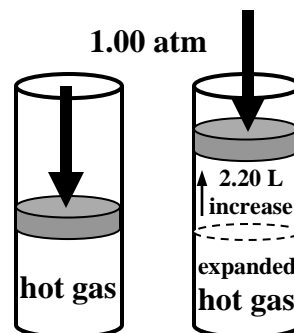
- Ordinarily, the  $P_{\text{external}}\Delta V$  value for a system is very small in comparison to its  $\Delta H$  value or its  $\Delta E$  value. Thus, for all practical purposes  $\Delta H \approx \Delta E$

**Q1:** Work is done by a gaseous system if the gases expand outwardly against the surroundings at a constant external pressure. How much work is done by the hot gaseous products of combustion in an engine if the piston chamber expands by 2.20 L at an external pressure of 1.00 atm? [ $1 \text{ L}\cdot\text{atm} = 101.3 \text{ J}$ ]

**A1:**  $w = -P_{\text{external}}\Delta V$

$$X = -(1.00 \text{ atm})(2.20 \text{ L}) = -2.20 \text{ L}\cdot\text{atm}$$

$$X = -2.20 \text{ L}\cdot\text{atm} \times \frac{101.3 \text{ J}}{1 \text{ L}\cdot\text{atm}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = \boxed{-0.223 \text{ kJ}}$$



**Q2:** The following reaction  $\text{CH}_{4(g)} + 2 \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2 \text{H}_2\text{O}_{(l)}$  ( $\Delta H = -890. \text{ kJ}$ ) occurs at constant external pressure and  $25^\circ\text{C}$ .  
 (a) Calculate the PV work for the reaction.  
 (b) Calculate the  $\Delta E$  for the reaction.

**A2:** (a)  $P_{\text{external}}\Delta V = \Delta n(RT)$  ...where  $\Delta n = \Delta \text{mol of gas}$

$$P_{\text{external}}\Delta V = (-2)(.08206)(298)$$

$$P_{\text{external}}\Delta V = -48.9 \text{ L}\cdot\text{atm}$$

$$P_{\text{external}}\Delta V = -48.9 \text{ L}\cdot\text{atm} \times \frac{101.3 \text{ J}}{1 \text{ L}\cdot\text{atm}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -4.95 \text{ kJ}$$

$$w = -P_{\text{external}}\Delta V = \boxed{+4.95 \text{ kJ}} \text{ work is done on the system!}$$

(b)  $\Delta H = \Delta E + P\Delta V$

$$-890. \text{ kJ} = X + (+4.95 \text{ kJ})$$

$$X = \boxed{-895 \text{ kJ}}$$

- In chemistry, we have (i) **electrical work** and (ii) **PV work**.

$$(i) w = -kQ_1Q_2 \left[ \frac{1}{d_f} - \frac{1}{d_i} \right]$$

$$(ii) w = -P_{\text{external}}\Delta V$$



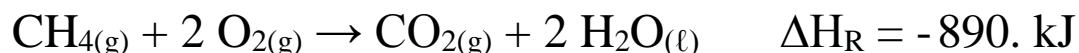
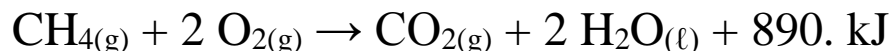
## IV – Enthalpy of Reaction

- The enthalpy change that accompanies a reaction is called the enthalpy of reaction or simply the heat of reaction. ( $\Delta H_R$ )

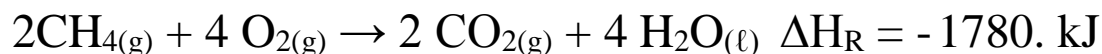
In endothermic reactions:  $\Delta H_R = (+)$

In exothermic reactions:  $\Delta H_R = (-)$

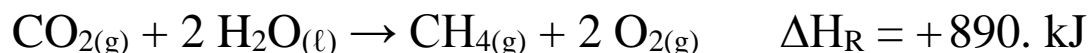
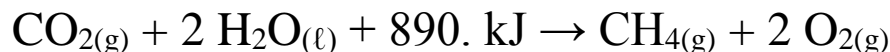
- A balanced chemical equation that shows the reaction's enthalpy change ( $\Delta H_R$ ) is called a thermochemical equation.



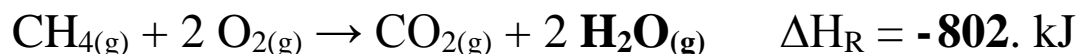
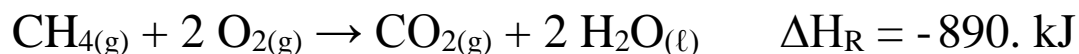
- Enthalpy change ( $\Delta H_R$ ) is an extensive property (*the amount matters*).



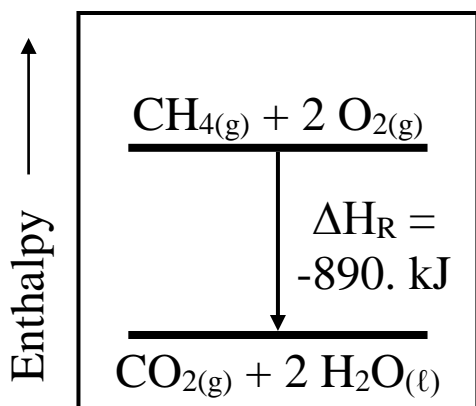
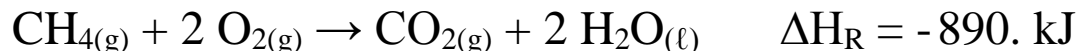
- The  $\Delta H_R$  for a reverse reaction, is equal in magnitude but opposite in sign.



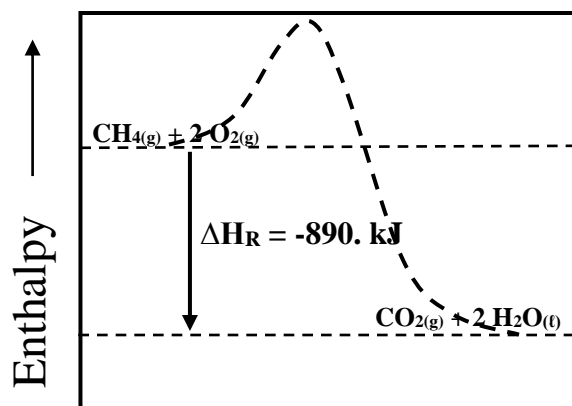
- The  $\Delta H_R$  for a reaction depends on the phase of the reactants and products.



- The  $\Delta H_R$  for a reaction may be represented by an enthalpy diagram.



or



**Q3:** How many kilojoules of heat are released if 305800 L of methane gas at 298 K and 1.00 atm is burned [ $d = 0.6556 \text{ g/L}$  at 298 K]?



$$305800 \text{ L} \times \frac{0.6556 \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{16.04 \text{ g}} \times \frac{-802 \text{ kJ}}{1 \text{ mol CH}_4} = -1.00 \times 10^7 \text{ kJ}$$

**Q4:** Complete combustion of 1.0 metric ton of coal (*assuming pure carbon*) to gaseous carbon dioxide releases  $3.3 \times 10^{10} \text{ J}$  of heat. How many metric tons of coal would provide the same energy output as determined in question-3?

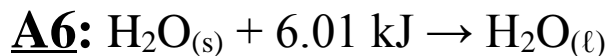
$$\text{A4: } -1.00 \times 10^7 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1.0 \text{ metric ton}}{-3.3 \times 10^{10} \text{ J}} = 0.30 \text{ metric ton coal}$$

**Q5:** How many kilojoules of heat are released when a cubic meter of water at  $0.0^\circ\text{C}$  freezes? [*heat of fusion of ice is  $+6.01 \text{ kJ/mol}$* ]



$$1 \text{ m}^3 \times \frac{(100 \text{ cm})^3}{(1 \text{ m})^3} \times \frac{1 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{-6.01 \text{ kJ}}{1 \text{ mol}} = -334000 \text{ kJ}$$

**Q6:** How many food Calories are burned to just melt one ice cube whose liquid water volume is  $18.0 \text{ cm}^3$ ? [ $1 \text{ Cal} = 4.184 \text{ kJ}$ ]



$$18.0 \text{ cm}^3 \times \frac{1 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{+6.01 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} = 1.43 \text{ Cal}$$

**Q7:** 1.0 lb of body fat is equivalent to 4100 Cal. How many of the ice cubes from question-6 would be needed to burn 1.0 lb of body fat?

$$\text{A7: } 4100 \text{ Cal} \times \frac{1 \text{ ice cube}}{1.43 \text{ Cal}} = 2900 \text{ ice cubes!}$$

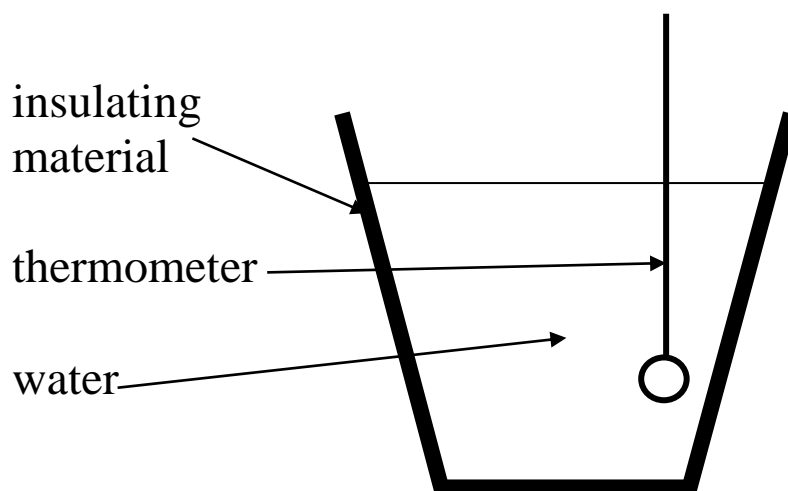
# Chapter-5

## Lesson-3

### V – Calorimetry

- The measurement of heat flow is called **calorimetry**.
- A **calorimeter** is a device used to measure heat flow.
- **Heat capacity** is the amount of heat a body needs to absorb to raise its temperature by 1°C (1 K) or the amount of heat a body needs to release when its temperature drops by 1°C.
- The heat capacity for 1 mole of a substance is called the **molar heat capacity**.
- The heat capacity for 1 g of a substance is called the **specific heat capacity** or just the **specific heat**.

#### A) Constant Pressure Calorimetry



- **Simple calorimeters** are not completely sealed. Thus the reactions that take place in them occur at constant pressure.
- Recall that at constant pressure  $\Delta H = q$ .
- The water in the calorimeter is considered as “the surroundings” and the material added to the calorimeter is considered as “the system”.
- Thus,  $\Delta H_R = - q_{\text{water}}$

- The heat absorbed or released by a substance as its temperature changes is determined by the equation:

$$q = mc\Delta T$$

where **q** is the heat absorbed or released (J), **m** is the mass of the substance (g), **c** is the specific heat of the substance (J/g°C), and  $\Delta T$  is the change in temperature of the substance (°C)

- The specific heat for water is 4.184 J/g°C.
- The specific heat for dilute solutions is approximately equal to 4.184 J/g°C.

**Q1:** A 20.6 g aluminum cylinder at 100.0°C is placed in 95.0 g of water at 21.0°C. The water's temperature rises to 24.5°C.

- Calculate the heat absorbed by the water.
- Calculate the specific heat of aluminum.

**A1:** (a)  $q_{\text{water}} = mc\Delta T$

$$q_{\text{water}} = (95.0 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(+3.5^\circ\text{C})$$

$$q_{\text{water}} = \boxed{+1400 \text{ J}}$$

Note:

*Both “q” and “ΔT” have signs.*

(b)  $c = q_{\text{aluminum}}/m\Delta T$

$$c = (-1400 \text{ J})/[(20.6 \text{ g})(-75.5^\circ\text{C})]$$

$$c = \boxed{0.90 \text{ J/g}^\circ\text{C}}$$

*-q = heat released*

*+q = heat absorbed*

$\Delta T = T_f - T_i$

**Q2:** 10.0 g LiBr dissolves in 250.0 g of water in a calorimeter. The water's temperature increases from 19.8°C to 25.3°C.

- What is the  $\Delta H_R$  (*in kJ/mol*) for the dissolving of LiBr?
- Is the dissolving reaction endothermic or exothermic?

**A2:** (a)  $q_{\text{water}} = mc\Delta T$

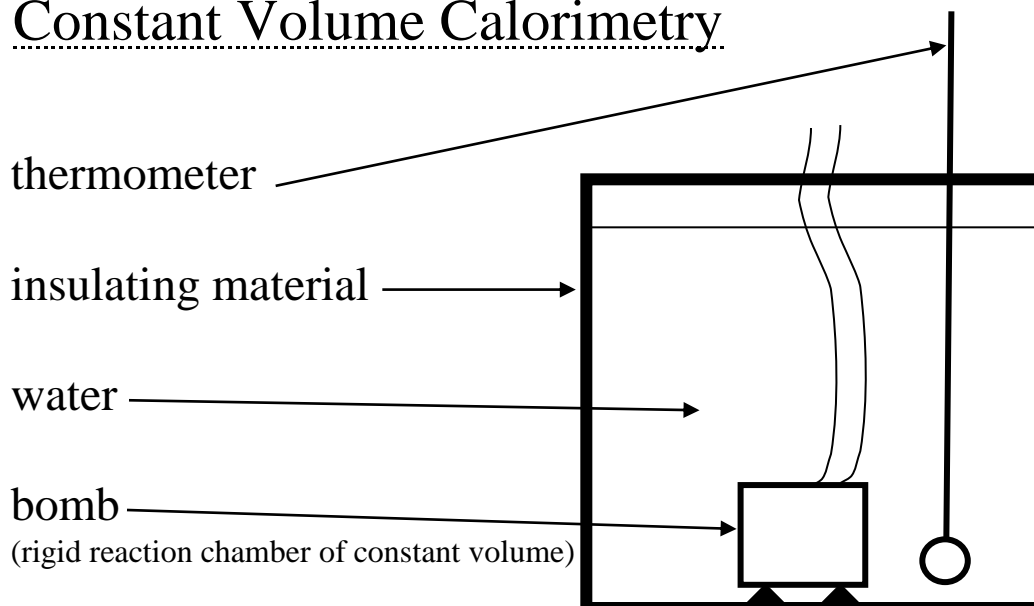
$$q_{\text{water}} = (250.0 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(+5.5^\circ\text{C}) = +5800 \text{ J} = +5.8 \text{ kJ}$$

$$q_{\text{LiBr}} = \frac{-5.8 \text{ kJ}}{10.0 \text{ g}} \times \frac{86.8 \text{ g}}{1 \text{ mol}} = -50.3 \text{ kJ/mol LiBr dissolved}$$

$$\Delta H_R = \boxed{-50.3 \text{ kJ/mol LiBr}}$$

(b)  $\boxed{\text{exothermic}}$

## B) Constant Volume Calorimetry



- In **bomb calorimeters**, the reaction chamber is designed to withstand high pressures at a constant volume.
- At constant volume,  $\Delta E = q_v$ . But recall that  $\Delta H \approx \Delta E$ .
- The amount of heat absorbed (or released) by the calorimeter is determined by the formula:

$$q_{\text{calorimeter}} = C_{\text{calorimeter}} \bullet \Delta T_{\text{water}}$$

where  $C_{\text{calorimeter}}$  is The **heat capacity of the calorimeter** or the simply the **calorimeter constant**.

- Sometimes, The amount of heat absorbed (or released) by the calorimeter is determined by the formula:

$$q_{\text{calorimeter}} = [(m_{\text{water}} \bullet c_{\text{water}}) + (B_{\text{bomb}})] \bullet \Delta T_{\text{water}}$$

where  $(m_{\text{water}} \bullet c_{\text{water}})$  is the heat capacity of the water in the calorimeter and  $B_{\text{bomb}}$  is the **heat capacity of the bomb** or the **bomb constant** of the calorimeter.

### **CAUTION:**

When reading a problem, determine whether you are given the value of the bomb constant or of the calorimeter constant for the calorimeter.

**Q3:** The combustion of 1.60 g of CH<sub>4</sub> in a bomb calorimeter with 3970.0 g of water at 21.77°C causes the water's temperature to increase to 26.93°C. What is the heat of combustion ( $\Delta H_C$ ) in **kJ/mol** if the bomb constant of the calorimeter is 526 J/°C?

**A3:**  $q_{\text{calorimeter}} = [(m_{\text{water}} \bullet c_{\text{water}}) + (B_{\text{bomb}})] \bullet \Delta T_{\text{water}}$

$$q_{\text{calorimeter}} = [(3970.0 \text{ g} \bullet 4.184 \text{ J/g}^\circ\text{C}) + (526 \text{ J/}^\circ\text{C})] \bullet (5.16^\circ\text{C})$$

$$q_{\text{calorimeter}} = \frac{88400 \text{ J}}{1.60 \text{ g}} \times \frac{1.000 \text{ kJ}}{1000 \text{ J}} \times \frac{16.0 \text{ g}}{1 \text{ mol}} = 884 \text{ kJ/mol CH}_4$$

$$\Delta H_C = - q_{\text{calorimeter}} = \boxed{- 884 \text{ kJ/mol CH}_4}$$

**Q4:** 3.16 g of salicylic acid (C<sub>7</sub>H<sub>6</sub>O<sub>3</sub>) is burned in a bomb calorimeter. The temperature of the water in the calorimeter rises by 2.83°C. What is the heat of combustion ( $\Delta H_C$ ) of the salicylic acid sample in **kJ/mole** if the calorimeter constant is 24500 J/°C?

**A4:**  $q_{\text{calorimeter}} = C_{\text{calorimeter}} \bullet \Delta T_{\text{water}}$

$$q_{\text{calorimeter}} = (24500 \text{ J/}^\circ\text{C}) \bullet (2.83^\circ\text{C}) = 69300 \text{ J}$$

$$q_{\text{calorimeter}} = \frac{69300 \text{ J}}{3.16 \text{ g}} \times \frac{1.000 \text{ kJ}}{1000 \text{ J}} \times \frac{138 \text{ g}}{1 \text{ mol}} = 3030 \text{ kJ/mol C}_7\text{H}_6\text{O}_3$$

$$\Delta H_C = - q_{\text{calorimeter}} = \boxed{- 3030 \text{ kJ/mol C}_7\text{H}_6\text{O}_3}$$

**Q5:** When 1.00 g of glucose was burned in a bomb calorimeter filled with 8060. g of water and the water's temperature rose by 4.30°C. It is known that 1.00 g of glucose releases 155660 J of heat when burned. What was the calorimeter's bomb constant in J/°C?

**A5:**  $q_{\text{calorimeter}} = [(m_{\text{water}} \bullet c_{\text{water}}) + (B_{\text{bomb}})] \bullet \Delta T_{\text{water}}$

$$155660 \text{ J} = [(8060. \text{ g} \bullet 4.184 \text{ J/g}^\circ\text{C}) + (X)] \bullet 4.30^\circ\text{C}$$

$$155660 \text{ J} = [(33720 \text{ J/}^\circ\text{C}) + (X)] \bullet 4.30^\circ\text{C}$$

$$\boxed{X = 2480 \text{ J/}^\circ\text{C}}$$

## VI – Hess's Law

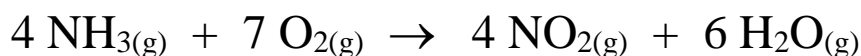
- The value of a reaction's enthalpy change ( $\Delta H$ ) is the same whether it occurs in one step or a series of steps. This is called **Hess' Law**.

**Q6:** What is the  $\Delta H_R$  for ...  $2 \text{ CO}_{(g)} + \text{O}_{2(g)} \rightarrow 2 \text{ CO}_{2(g)}$

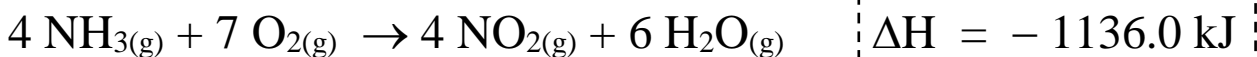
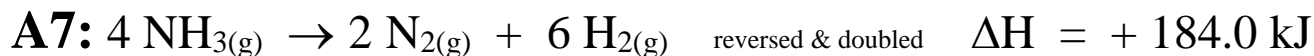
**Given:**



**Q7:** What is the  $\Delta H_R$  for ...

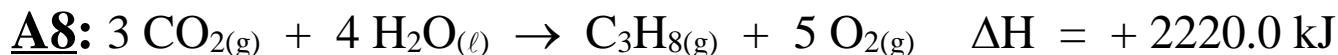
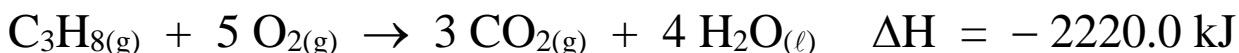


**Given:**

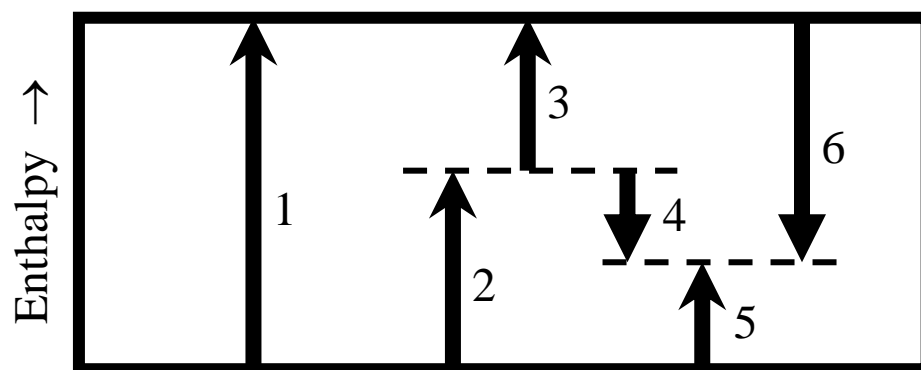


**Q8:** What is the  $\Delta H_R$  for ...  $3 \text{ C}_{(s)} + 4 \text{ H}_{2(g)} \rightarrow \text{C}_3\text{H}_{8(g)}$

**Given:**



**Q9:** Consider the energy diagram below.



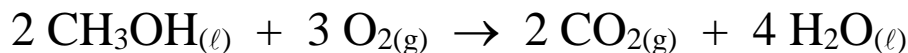
- Write an equation showing how  $\Delta H_1$  is related to  $\Delta H_2$  and  $\Delta H_3$ .
- Write an equation showing how  $\Delta H_5$  is related to  $\Delta H_1$  and  $\Delta H_6$ .
- Write an equation showing how  $\Delta H_4$  is related to any other two  $\Delta H$  values.

**A9:** (a)  $\Delta H_1 = \Delta H_2 + \Delta H_3$

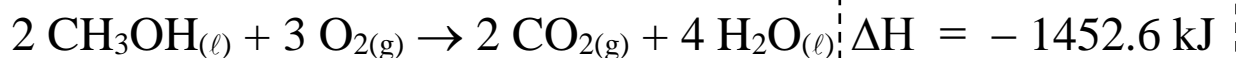
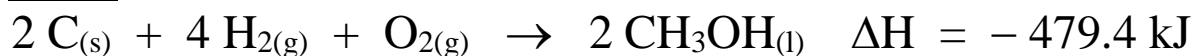
(b)  $\Delta H_5 = \Delta H_1 + \Delta H_6$

(c)  $\Delta H_4 = -\Delta H_2 + \Delta H_5$       or       $\Delta H_4 = \Delta H_6 + \Delta H_3$

**Q10:** What is the  $\Delta H_R$  for ...



**Given:**





# Chapter-5

## Lesson-4

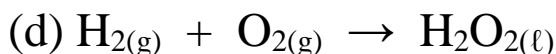
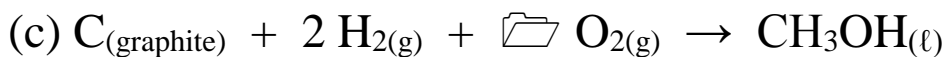
### VII – Enthalpy of Formation

- The heat absorbed or released when 1 mole of a compound is formed from its pure elements is called the **enthalpy (or heat) of formation** ( $\Delta H_f$ ).
- Since the amount of heat absorbed or released by a reaction is affected by the temperature of the system, the pressure of the system, and the physical states of the reactants and products, a set of standard conditions are used to standardizes all enthalpy values.
- The standard state of a substance (the phase at 298 K & 1 atm), 298 K, 1 atm (for gases), and 1 M (for aqueous solutions) are the set of standard conditions.
- The enthalpy change of a reaction at standard conditions is called the **standard enthalpy change** ( $\Delta H^\circ$ ).
- The compound's enthalpy of formation at standard conditions is called the **standard enthalpy of formation** ( $\Delta H_f^\circ$ ).
- The **standard enthalpy of formation** ( $\Delta H_f^\circ$ ) for any pure element in its standard state is zero.
- The **standard enthalpy of formation** ( $\Delta H_f^\circ$ ) for 1 M aqueous hydrogen ion ( $H^{1+}_{(aq)}$ ) is zero.

*See Appendix C in your textbook (page 1123) for  $\Delta H_f^\circ$  values.*

- Changes in enthalpy ( $\Delta H$ ) are not strongly temperature, pressure, or concentration dependent.
- Thus  $\Delta H_f^\circ$  values can be used for reactions that are not occurring at 298 K, 1 atm (for gases) or 1 M (for aqueous solutions).

**Q1:** Write the balanced equation for standard formation reaction for each of the following compounds (include phases).



- An enthalpy of a reaction ( $\Delta H_R$ ) or a heat of formation ( $\Delta H_f$ ) can be determined from the following form of Hess's Law.

$$\Delta H_{\text{Reaction}} = \sum \Delta H_f \text{ products} - \sum \Delta H_f \text{ reactants}$$

**Q2:** Using Appendix C, calculate the heat of reaction of



$$\begin{aligned} \mathbf{A2:} \Delta H &= [2(\Delta H_f \text{Li}^{1+}) + 2(\Delta H_f \text{OH}^{1-}) + 1(\Delta H_f \text{H}_2)] \\ &\quad - [2(\Delta H_f \text{Li}) + 2(\Delta H_f \text{H}_2\text{O}_{(\ell)})] \end{aligned}$$

$$\Delta H = [2(-278.5) + 2(-230.0) + (0)] - [2(0) + 2(-285.83)]$$

$$\Delta H = [-1017.0] - [-571.66] = \boxed{-445.3 \text{ kJ}} \quad (\text{exothermic})$$

**Q3:** Using Appendix C, calculate the heat of reaction of



$$\mathbf{A3:} \Delta H = [1(\Delta H_f \text{Na}^+) + 1(\Delta H_f \text{Cl}^{1-})] - [1(\Delta H_f \text{NaCl})]$$

$$\Delta H = [1(-240.1) + 1(-167.2)] - [1(-410.9)]$$

$$\Delta H = [-407.1] - [-410.9] = \boxed{+3.8 \text{ kJ}} \quad (\text{endothermic})$$

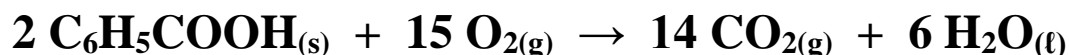
$$\Delta H_f \text{ of } \text{CH}_4(\text{g}) = -74.8 \text{ kJ/mol}$$

$$\Delta H_f \text{ of } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f \text{ of } \text{H}_2\text{O}(\text{g}) = -241.82 \text{ kJ/mol}$$

$$\Delta H_f \text{ of } \text{H}_2\text{O}(\text{l}) = -285.83 \text{ kJ/mol}$$

**Q4:** The combustion of 1 mol of benzoic acid releases 3222 kJ of heat.



Calculate the heat of formation ( $\Delta H_f$ ) for benzoic acid in kJ/mol.

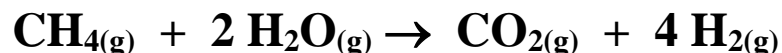
$$\textbf{A4: } \Delta H_C = [14(\Delta H_f \text{ CO}_2) + 6(\Delta H_f \text{ H}_2\text{O}(\text{l}))] - [2(\Delta H_f \text{ HC}_7\text{H}_5\text{O}_2) + 15(\Delta H_f \text{ O}_2)]$$

$$\underline{2}(-3222) = [14(-393.5) + 6(-285.83)] - [2(X) + 15(0)]$$

$$\underline{2}X = [-7224.0] - [-6444] \rightarrow \underline{2}X = -780 \text{ kJ} \quad (\textit{exothermic})$$

$$\boxed{\Delta H_f = -390 \text{ kJ/mol}}$$

**Q5:** Calculate the heat of reaction when methane reacts with steam.



$$\textbf{A5: } \Delta H = [1(\Delta H_f \text{ CO}_2) + 4(\Delta H_f \text{ H}_2)] - [1(\Delta H_f \text{ CH}_4) + 2(\Delta H_f \text{ H}_2\text{O}(\text{g}))]$$

$$\Delta H = [1(-393.5) + 4(0)] - [1(-74.8) + 2(-241.82)]$$

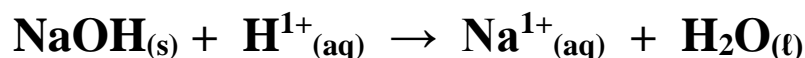
$$\Delta H = [-393.5] - [-558.44] = \boxed{+164.9 \text{ kJ}} \quad (\textit{endothermic})$$

$$\Delta H_f \text{ of } \text{Na}^{1+}_{(\text{aq})} = - 240.1 \text{ kJ/mol}$$

$$\Delta H_f \text{ of } \text{NaOH}_{(\text{s})} = - 425.6 \text{ kJ/mol}$$

$$\Delta H_f \text{ of } \text{OH}^{1-}_{(\text{aq})} = - 230.0 \text{ kJ/mol}$$

**Q6:** Calculate the heat of reaction when solid sodium hydroxide is neutralized by a strong acid.



$$\textbf{A6: } \Delta H = [1(\Delta H_f \text{Na}^+) + 1(\Delta H_f \text{H}_2\text{O}_{(\text{l})})] - [1(\Delta H_f \text{NaOH}) + 1(\Delta H_f \text{H}^+)]$$

$$\Delta H = [1(-240.1) + 1(-285.83)] - [(-425.6) + 1(0)]$$

$$\Delta H = [-525.9] - [-425.6] = \boxed{- 100.3 \text{ kJ}} \text{ (} \textit{exothermic} \text{)}$$

**Q7:** Calculate the heat of reaction when solid sodium metal is added to water.



$$\textbf{A7: } \Delta H = [2(\Delta H_f \text{Na}^+) + 2(\Delta H_f \text{OH}^-) + 1(\Delta H_f \text{H}_2)] - [2(\Delta H_f \text{Na}) + 2(\Delta H_f \text{H}_2\text{O}_{(\text{l})})]$$

$$\Delta H = [2(-240.1) + 2(-230.0) + 1(0)] - [2(0) + 2(-285.83)]$$

$$\Delta H = [-940.2] - [-571.66] = \boxed{- 368.5 \text{ kJ}} \text{ (} \textit{exothermic} \text{)}$$